

REMARKS

1. Restriction

Restriction has been required to one of five identified groups of claims.

- I. Claims 1–13
- II. Claims 14–19
- III. Claims 20–28
- IV. Claims 29–34
- V. Claims 35–37

Applicants affirm the election to prosecute the claims of Group III without traverse. Claims 1–19 and 29–37 have, accordingly, been canceled.

2. Examination

Claims 20–28 have been examined. Claims 24 and 25 stand rejected under the first paragraph of 35 U.S.C. § 112; Claims 20–22 stand rejected under 35 U.S.C. § 103(a) as unpatentable over U.S. Patent No. 5,520,682 (“Baust”) in view of U.S. Patent No. 3,889,680 (“Armao”) and further in view of U.S. Patent No. 5,147,538 (“Wright”) and even further in view of U.S. Patent No. 6,848,502 (“Bishop”); Claim 23 stands rejected under 35 U.S.C. § 103(a) as unpatentable over Baust in view of Armao and further in view of Wright and even further in view of Bishop and even further in view of U.S. Patent No. 5,275,595 (“Dobak”); Claims 26 and 27 stand rejected under 35 U.S.C. § 103(a) as unpatentable over Baust in view of Armao and further in view of Wright and even further in view of Bishop and even further in view of U.S. Patent No. 4,519,389 (“Gudkin”); and Claim 28 stands rejected under 35 U.S.C. § 103(a) as unpatentable over Baust in view of Armao and further in view of Wright and even further in view of Bishop and even further in view of U.S. Patent No. 5,741,248 (“Stern”).

a. § 112 Rejections

The § 112 rejections are respectfully traversed. The Office Action indicates that there is no support provided in the specification to enable one of skill in the art to cool an imaging array or electronic circuit using the methods described therein, and that “[t]he premise of the instant application is toward a cryoprobe for treating biological tissue” (Office Action, p. 4). While the written description does provide a number of specific illustrations in which biological tissue is treated, it also notes the general applicability of the near-critical cooling technique that is described. The Examiner’s attention is drawn to the following disclosure in the specification.

While this improved method of cooling is described herein for application to probes for cryosurgery (see discussion of Fig. 7 below), it has very broad applicability to all cryogenic cooling applications involving a flow of liquid cryogen. More generally, the methods described herein may be used to cool other devices with any cryogen flow operating near its critical point. These devices may include sensors within image plane arrays for remote sensing applications, superconducting wires and cables, and all other devices requiring cryogenic support. Other cryogens of interest using this technique include and liquefied gas, including argon, neon, xenon, helium, hydrogen, and oxygen, to name a few. These are single component flowing liquid-vapor systems. Mixtures of substances in flowing liquid vapor systems may also be used in this way, such as flowing hydrocarbon gas mixtures, provided that their critical point pressures and temperatures are readily obtained within the engineered cooling or liquid-gas processing system.

(Application, p. 20, l. 30 – p. 21, l. 9).

In addition to noting this general applicability for cryogenic cooling, large sections of the disclosure are not in any way limited to the specific applications of treating biological tissue. For example, the disclosure at p. 7, l. 27 – p. 16, l. 26 provides a general description of the cooling techniques and provides a comparison with certain prior art techniques such as Joule-Thomson cooling. The disclosure at p. 17, l. 1 – p. 25, l. 13 provides a general description of specific structures in which the methods for cooling may be embodied; while this disclosure does sometimes make reference to biological applications and describes certain

features of the structures that may be of particular interest in biological applications, the overall description of those structures is more general.

It is, accordingly, believed that Claims 24 and 25 are fully supported by an enabling disclosure.

b. § 103 Rejections

Independent Claim 20 has been amended to clarify the nature of the “critical point.” As used throughout the application, in both the specification and claims, references to a “critical point” are intended to refer to a specific physical property of substances. This property may be understood with reference to Fig. 1 of the Application, which shows a phase diagram — the numerical values of the pressure and temperature axes are suitable for nitrogen, but the general shape of the diagram is typical of many substances. The phase diagram shows that under different pressure and temperature states, the substance may be in different physical phases, with the diagram identifying regions where it is in a “gas” phase, a “liquid” phase, or a “solid” phase. A common characteristic of substances is that the line in this pressure-temperature space that separates the liquid and gas states terminates in a point (shown in Fig. 1 but not identified with a reference number). This point is commonly referred to in the art as the “critical point” and has a temperature and pressure characteristic of the substance. *See also* Application, p. 9, ll. 11 – 16.

To illustrate that the term is well known in the art and refers unambiguously to this characteristic, general material describing phase diagrams is provided as Exhibit 1. A discussion of critical points (identified in the phase diagrams in this material by the label “C”) begins at p. 7.

There are a number of different ways in which the critical point may be characterized physically. Independent Claim 20 has been amended to include language defining the critical point as where molar volumes are substantially equivalent for liquid and gas phases (*see* Application, p. 9, ll. 7–8). This merely makes explicit what was previously implicit. For

purposes of clarification, the cryogen is also now referred to as a “fluid,” with corresponding amendments being made to Claims 23 and 27.

It is respectfully noted with these clarifications that the cited art does not teach or suggest circulation of a cryogenic fluid under physical conditions near the critical point. The application explains at length how such circulation results in the avoidance of vapor lock, a phenomenon that the cited art does recognize as an undesirable consequence of prior art approaches. Such prior art approaches use *evaporative cooling* techniques, which characteristically operate well away from the critical point. This is particularly evident in Joule-Thomson approaches where rapid expansion and decompression of high-pressure *gases* produce cooling from the Joule-Thomson effect at a probe tip. Other prior art techniques use flows of *liquid* cryogens that create cooling during evaporation at, or near, a probe tip. The fact that these prior art approaches use substances having a well-defined phase of “liquid” or “gas” clearly distinguishes them from the claimed approach since such distinctions do not exist near the critical point.

The Office Action cites the following language in Baust in particular for its disclosure of a “critical point” (Office Action, pp. 4–5).

With regard to total nitrogen consumption as a function of hole pattern (size, number and position) it is noted that for two-phase flow (liquid-gas) the discharge rate to the probe tip increases uniformly as the supply pressure increases, at least in the lower supply pressure range. Gradually, a critical point will be reached where the mass flow of nitrogen escaping from the discharge end of the supply tube achieves a maximum, which then cannot be exceeded by further increasing the supply pressure.

The prediction of how the critical point is complicated by the nature of cryogenic flow, largely because of the changes in vapor and liquid ratio (quantity) and the mass, heat and momentum exchanges taking place at the vapor-liquid interface. The critical point of mass flow, however, can be adequately estimated by an extended experiment in which a homogenous flow to the probe tip is maintained by sub-cooling the LN₂ and subsequently minimizing the heat loss along the LN₂ transport system.

Since the pressure decreases as nitrogen flows in the supply tube, the rate of nitrogen gas or liquid escaping from the hose will vary from hole to hole, depending upon the position of the hole on the supply tube. At maximum operating conditions, however, the mass flow m of LN_2 discharged to the probe tip will be a constant, regardless of the hole pattern. Here, the holes are assumed to function only to vent the gas and accelerate the process to attain the maximum flow conditions.

(Baust, Col. 13, ll. 16 – 38).

It is clear from this passage that what Baust identifies as a “critical point” is different from the critical point of the phase diagram. Indeed, it is a “critical point of mass flow” (emphasis added). Baust is, in fact, detailing the problems of *evaporative cooling*, and maintaining flow from the boiling of nitrogen and the pressure increases at the liquid-vapor interface. This actually teaches away from the claim limitation by describing operation in a completely different region of the phase diagram and illustrates the disadvantages of the prior art that are avoided by operating in a region of the phase diagram where there is no well defined liquid or gas.

Since the claim limitation is not taught or suggested by the prior art, independent Claim 20 is believed to be patentable. Each of the dependent claims is also believed to be patentable by virtue of its dependence from a patentable claim.

It is noted, however, that the Office Action cites Dobak as disclosing a certain pressure range. The following additional comments regarding Dobak are, accordingly, offered in the interest of advancing prosecution of the application.

To make out a *prima facie* case under § 103(a) to combine Dobak with the other cited art, all of the claimed limitations must be identified in the cited art and a motivation for their combination must be identified. In this instance, it is respectfully noted that notwithstanding the assertion in the Office Action, Dobak does not teach or suggest the limitation of Claim 23 embodying the use of liquid nitrogen at a pressure of about 33.5 atm. The Office Action cites the disclosure of Claim 9, but this disclosure is concerned with “using mixed


gas” (emphasis added). Such mixtures are central to the disclosure of Dobak (*see, e.g.*, the table in Col. 4), which otherwise uses a conventional Joule-Thomson approach based on evaporative cooling (*id.*, Col. 3, l. 33 – Col. 4, l. 10). This distinction is also relevant to the lack of a motivation to combine Dobak with Baust. One of skill in the art would have no reason to use a physical operating range determined to be suitable for a gas mixture with the liquid-nitrogen arrangement described in Baust because the thermodynamic properties of the substances are consequently different. Furthermore, even if a *prima facie* case could be established for the combination, the claimed conditions exhibit criticality as illustrated by evidence provided in the application of improved cooling efficiency and avoidance of vapor lock (Application, p. 11, l. 15 – p. 16, l. 26). *See* MPEP 2144.05.

3. Conclusion

In view of the foregoing, Applicants believe all claims now pending in this application are in condition for allowance. The issuance of a formal Notice of Allowance at an early date is respectfully requested.

If the Examiner believes a telephone conference would expedite prosecution of this application, please telephone the undersigned at 303-571-4000.

Respectfully submitted,


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Date: November 23, 2005

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PHASE DIAGRAMS OF PURE SUBSTANCES

This page explains how to interpret the phase diagrams for simple pure substances - including a look at the special cases of the phase diagrams of water and carbon dioxide. This is going to be a long page, because I have tried to do the whole thing as gently as possible.

The basic phase diagram

What is a phase?

At its simplest, a phase can be just another term for solid, liquid or gas. If you have some ice floating in water, you have a solid phase present and a liquid phase. If there is air above the mixture, then that is another phase.

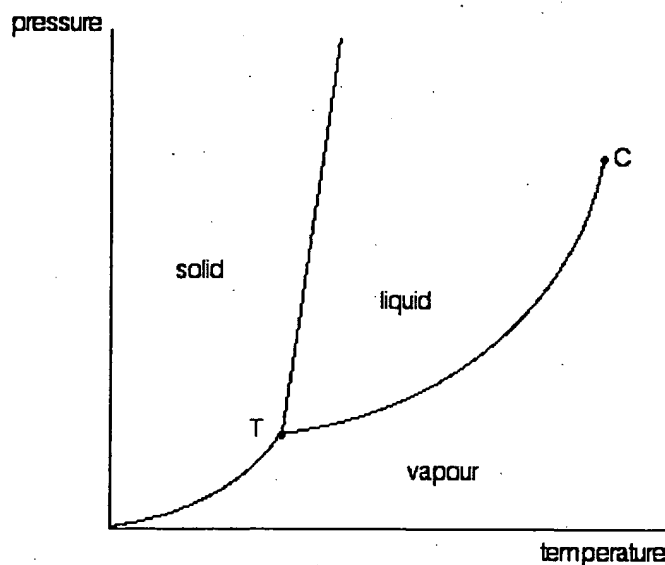
But the term can be used more generally than this. For example, oil floating on water also consists of two phases - in this case, two liquid phases. If the oil and water are contained in a bucket, then the solid bucket is yet another phase. In fact, there might be more than one solid phase if the handle is attached separately to the bucket rather than moulded as a part of the bucket.

You can recognise the presence of the different phases because there is an obvious boundary between them - a boundary between the solid ice and the liquid water, for example, or the boundary between the two liquids.

Phase diagrams

A phase diagram lets you work out exactly what phases are present at any given temperature and pressure. In the cases we'll be looking at on this page, the phases will simply be the solid, liquid or vapour (gas) states of a pure substance.

This is the phase diagram for a typical pure substance.

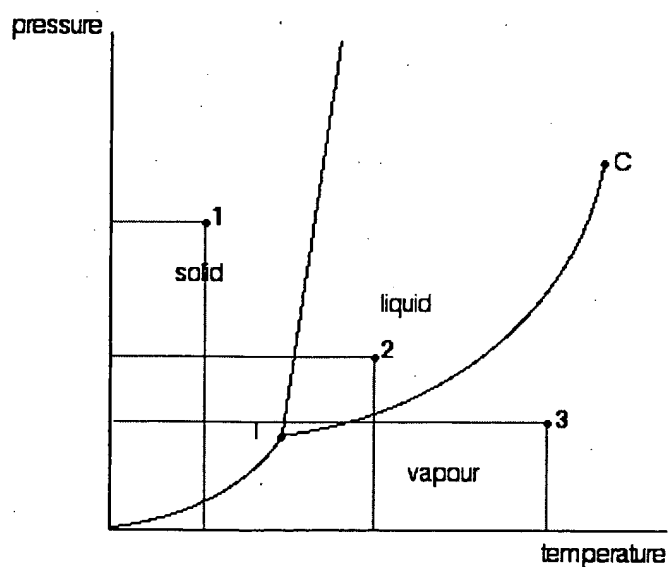


These diagrams (including this one) are nearly always drawn highly distorted in order to see what is going on more easily. There are usually two major distortions. We'll discuss these when they become relevant.

If you look at the diagram, you will see that there are three lines, three areas marked "solid", "liquid" and "vapour", and two special points marked "C" and "T".

The three areas

These are easy! Suppose you have a pure substance at three different sets of conditions of temperature and pressure corresponding to 1, 2 and 3 in the next diagram.



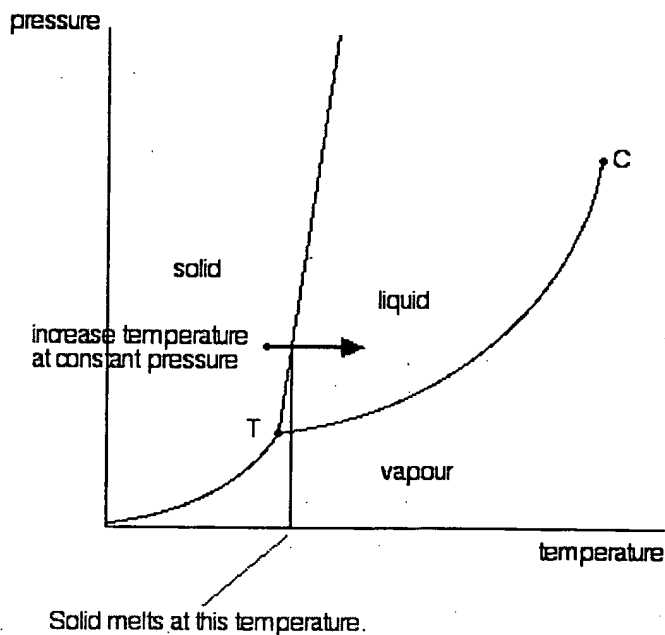
Under the set of conditions at 1 in the diagram, the substance

would be a solid because it falls into that area of the phase diagram. At 2, it would be a liquid; and at 3, it would be a vapour (a gas).

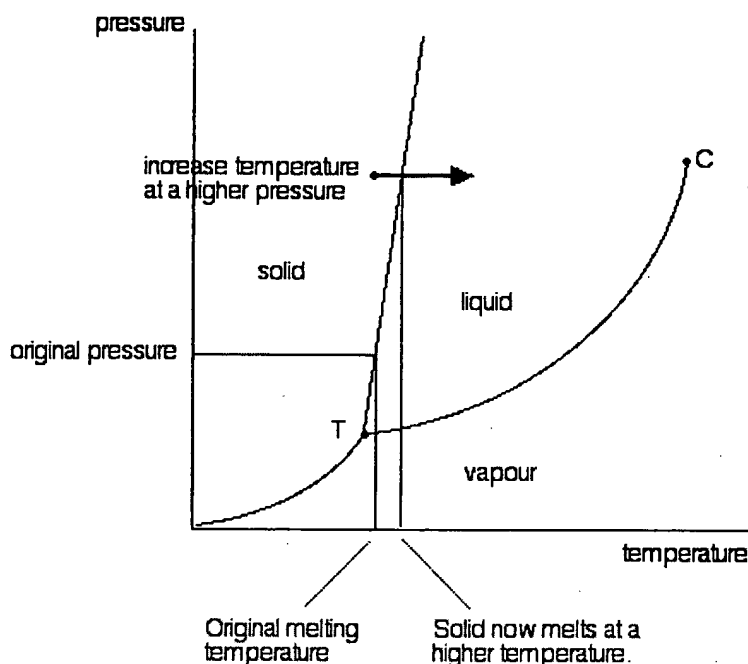
Note: I'm using the terms vapour and gas as if they were interchangeable. There are subtle differences between them that I'm not ready to explain for a while yet. Be patient!

Moving from solid to liquid by changing the temperature:

Suppose you had a solid and increased the temperature while keeping the pressure constant - as shown in the next diagram. As the temperature increases to the point where it crosses the line, the solid will turn to liquid. In other words, it melts.



If you repeated this at a higher fixed pressure, the melting temperature would be higher because the line between the solid and liquid areas slopes slightly forward.



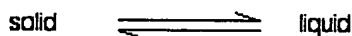
Note: This is one of the cases where we distort these diagrams to make them easier to discuss. This line is much more vertical in practice than we normally draw it. There would be very little change in melting point at a higher pressure. The diagram would be very difficult to follow if we didn't exaggerate it a bit.

So what actually is this line separating the solid and liquid areas of the diagram?

It simply shows the effect of pressure on melting point.

Anywhere on this line, there is an equilibrium between solid and liquid.

You can apply Le Chatelier's Principle to this equilibrium just as if it was a chemical equilibrium. If you increase the pressure, the equilibrium will move in such a way as to counter the change you have just made.



For most substances, the solid occupies a smaller volume than the liquid.

If it converted from liquid to solid, the pressure would tend to decrease again because the solid takes up slightly less space for most substances.

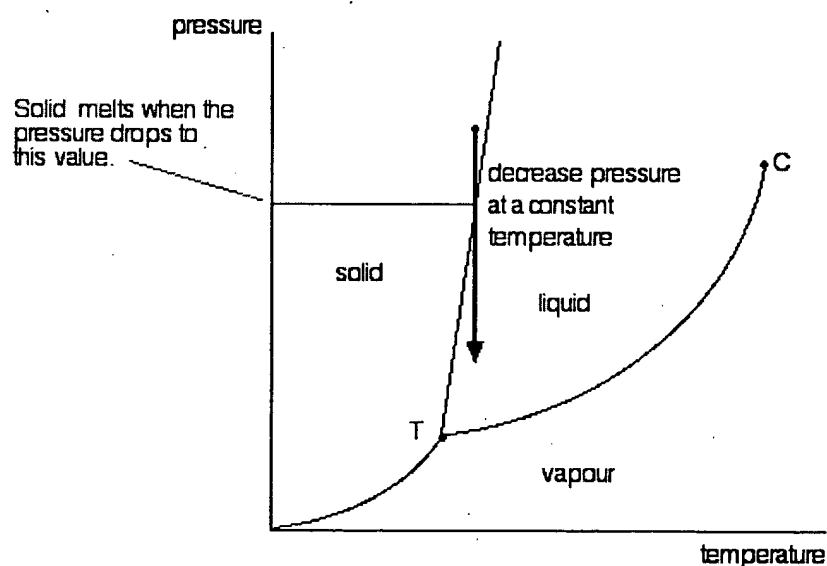
That means that increasing the pressure on the equilibrium mixture of solid and liquid at its original melting point will convert the mixture back into the solid again. In other words, it will no longer melt at this temperature.

To make it melt at this higher pressure, you will have to increase the temperature a bit. Raising the pressure raises the melting point of most solids. That's why the melting point line slopes forward for most substances.

Moving from solid to liquid by changing the pressure:

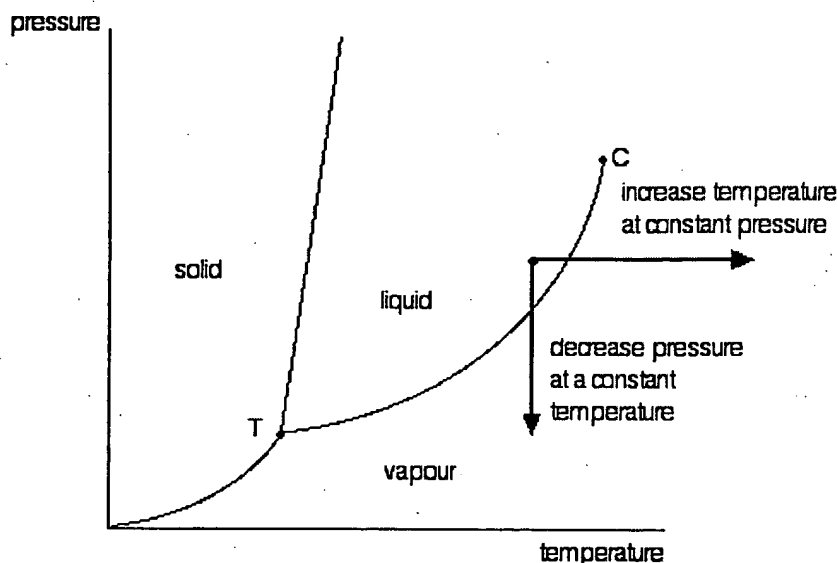
You can also play around with this by looking at what happens if you decrease the pressure on a solid at constant temperature.

Note: You have got to be a bit careful about this, because exactly what happens if you decrease the pressure depends on exactly what your starting conditions are. We'll talk some more about this when we look at the line separating the solid region from the vapour region.



Moving from liquid to vapour:

In the same sort of way, you can do this either by changing the temperature or the pressure.



The liquid will change to a vapour - it boils - when it crosses the boundary line between the two areas. If it is temperature that you are varying, you can easily read off the boiling temperature from the phase diagram. In the diagram above, it is the temperature where the red arrow crosses the boundary line.

So, again, what is the significance of this line separating the two areas?

Anywhere along this line, there will be an equilibrium between the liquid and the vapour. The line is most easily seen as the effect of pressure on the boiling point of the liquid.

As the pressure increases, so the boiling point increases.

Note: I don't want to make any very big deal over this, but this line is actually exactly the same as the graph for the effect of temperature on the saturated vapour pressure of the liquid. Saturated vapour pressure is dealt with on a separate page. A liquid will boil when its saturated vapour pressure is equal to the external pressure.

Suppose you measured the saturated vapour pressure of a liquid at 50°C, and it turned out to be 75 kPa. You could plot that as one point on a vapour pressure curve, and then go on to measure other saturated vapour pressures at different temperatures and plot those as well.

Now, suppose that you had the liquid exposed to a total external pressure of 75 kPa, and gradually increased the temperature. The liquid would boil when its saturated vapour pressure became equal to the external pressure - in this case at 50°C. If you have the complete vapour pressure curve, you could equally well find the boiling point corresponding to any other external pressure.

That means that the plot of saturated vapour pressure against temperature is exactly the same as the curve relating boiling point and

external pressure - they are just two ways of looking at the same thing.

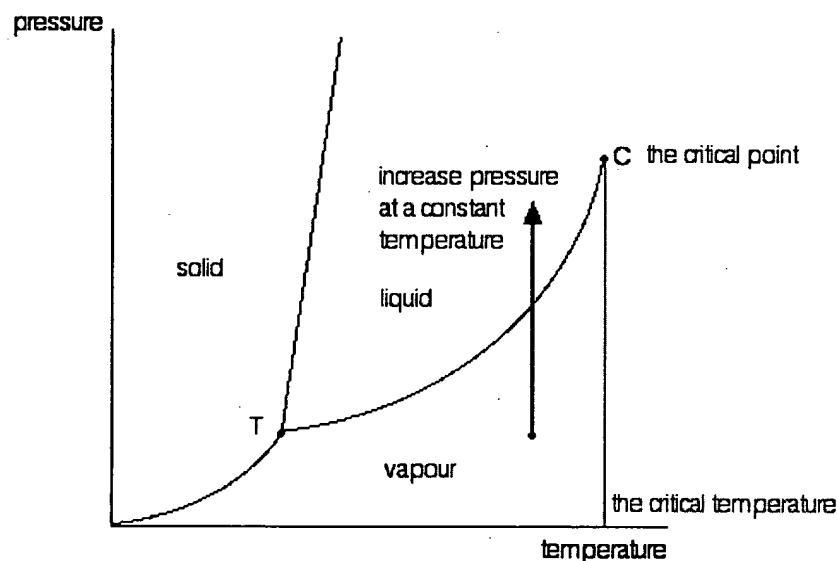
If all you are interested in doing is interpreting one of these phase diagrams, you probably don't have to worry too much about this.

The critical point

You will have noticed that this liquid-vapour equilibrium curve has a top limit that I have labelled as **C** in the phase diagram.

This is known as the **critical point**. The temperature and pressure corresponding to this are known as the **critical temperature** and **critical pressure**.

If you increase the pressure on a gas (vapour) at a temperature lower than the critical temperature, you will eventually cross the liquid-vapour equilibrium line and the vapour will condense to give a liquid.



This works fine as long as the gas is below the critical temperature. What, though, if your temperature was *above* the critical temperature? There wouldn't be any line to cross!

That is because, above the critical temperature, it is impossible to condense a gas into a liquid just by increasing the pressure. All you get is a highly compressed gas. The particles have too much energy for the intermolecular attractions to hold them together as a liquid.

The critical temperature obviously varies from substance to substance and depends on the strength of the attractions

between the particles. The stronger the intermolecular attractions, the higher the critical temperature.

Note: This is now a good point for a quick comment about the use of the words "gas" and "vapour". To a large extent you just use the term which feels right. You don't usually talk about "ethanol gas", although you *would* say "ethanol vapour". Equally, you wouldn't talk about oxygen as being a vapour - you always call it a gas.

There are various guide-lines that you can use if you want to. For example, if the substance is commonly a liquid at or around room temperature, you tend to call what comes away from it a vapour. A slightly wider use would be to call it a vapour if the substance is below its critical point, and a gas if it is above it. Certainly it would be unusual to call anything a vapour if it was above its critical point at room temperature - oxygen or nitrogen or hydrogen, for example. These would all be described as gases.

This is absolutely NOT something that is at all worth getting worked up about!

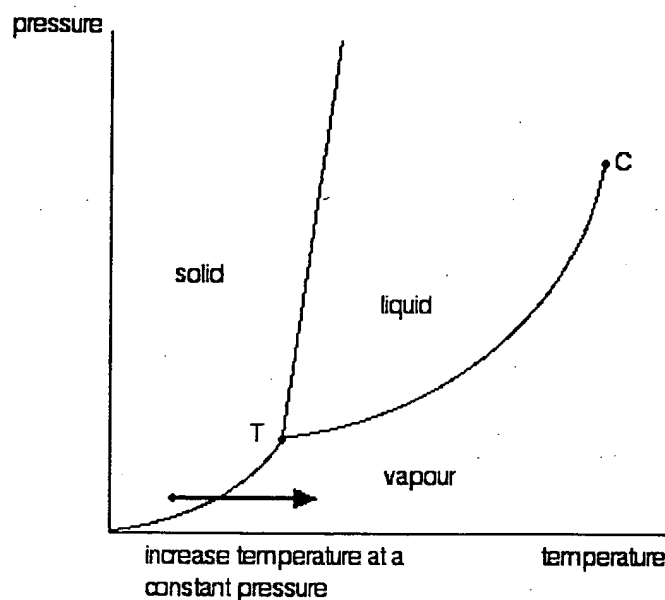
Moving from solid to vapour:

There's just one more line to look at on the phase diagram. This is the line in the bottom left-hand corner between the solid and vapour areas.

That line represents solid-vapour equilibrium. If the conditions of temperature and pressure fell exactly on that line, there would be solid and vapour in equilibrium with each other - the solid would be subliming. (Sublimation is the change directly from solid to vapour or vice versa without going through the liquid phase.)

Once again, you can cross that line by either increasing the temperature of the solid, or decreasing the pressure.

The diagram shows the effect of increasing the temperature of a solid at a (probably very low) constant pressure. The pressure obviously has to be low enough that a liquid can't form - in other words, it has to happen below the point labelled as T.



You could read the sublimation temperature off the diagram. It will be the temperature at which the line is crossed.

The triple point

Point T on the diagram is called the ***triple point***.

If you think about the three lines which meet at that point, they represent conditions of:

- solid-liquid equilibrium
- liquid-vapour equilibrium
- solid-vapour equilibrium

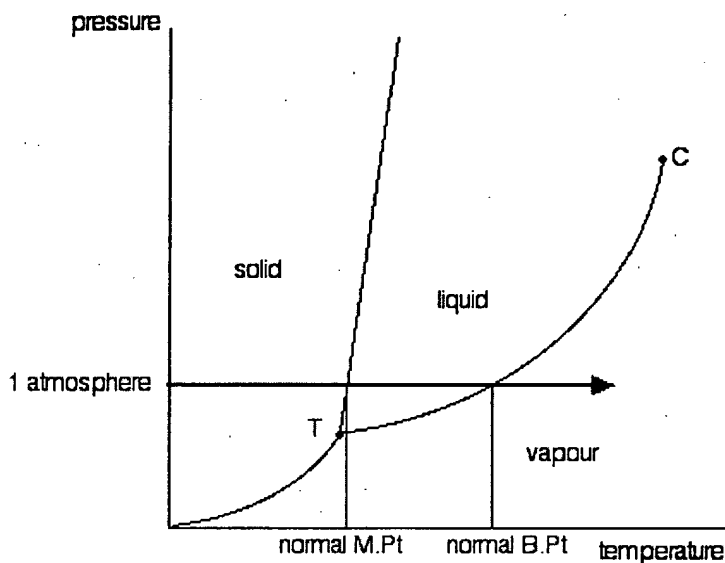
Where all three lines meet, you must have a unique combination of temperature and pressure where all three phases are in equilibrium together. That's why it is called a ***triple point***.

If you controlled the conditions of temperature and pressure in order to land on this point, you would see an equilibrium which involved the solid melting and subliming, and the liquid in contact with it boiling to produce a vapour - and all the reverse changes happening as well.

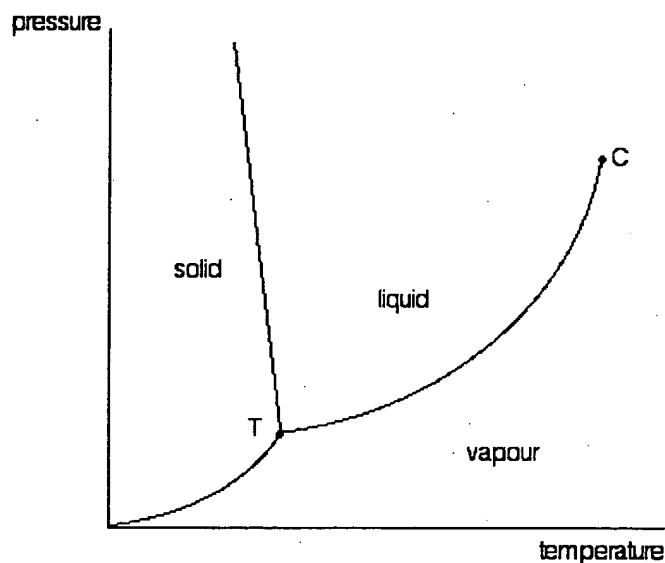
If you held the temperature and pressure at those values, and kept the system closed so that nothing escaped, that's how it would stay. A strange set of affairs!

Normal melting and boiling points

The normal melting and boiling points are those when the pressure is 1 atmosphere. These can be found from the phase diagram by drawing a line across at 1 atmosphere pressure.

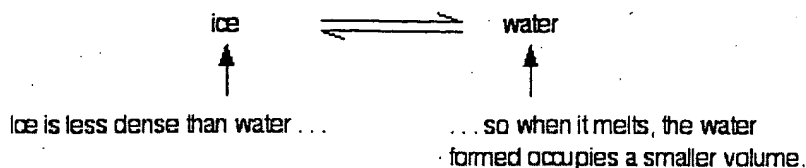


The phase diagram for water



There is only one difference between this and the phase diagram that we've looked at up to now. The solid-liquid equilibrium line (the melting point line) slopes backwards rather than forwards.

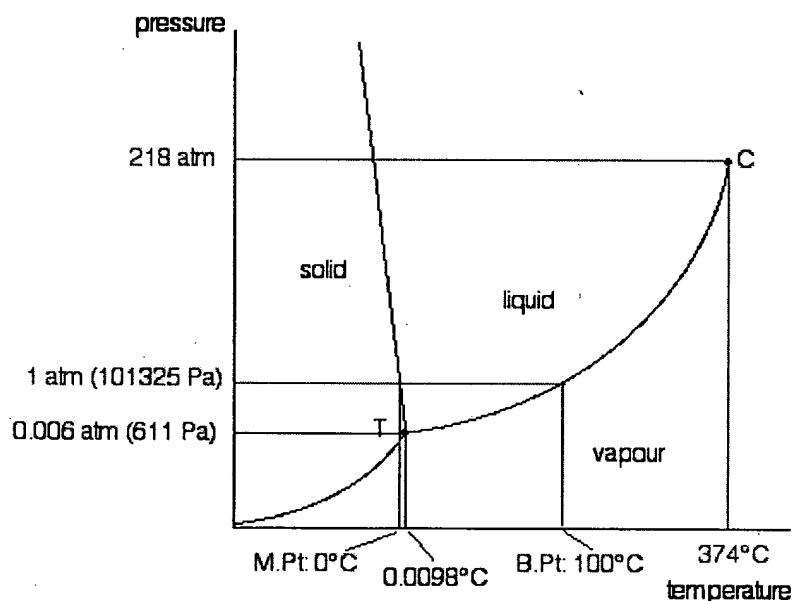
In the case of water, the melting point gets lower at higher pressures. Why?



If you have this equilibrium and increase the pressure on it, according to Le Chatelier's Principle the equilibrium will move to reduce the pressure again. That means that it will move to the side with the smaller volume. Liquid water is produced.

To make the liquid water freeze again at this higher pressure, you will have to reduce the temperature. Higher pressures mean lower melting (freezing) points.

Now lets put some numbers on the diagram to show the exact positions of the critical point and triple point for water.



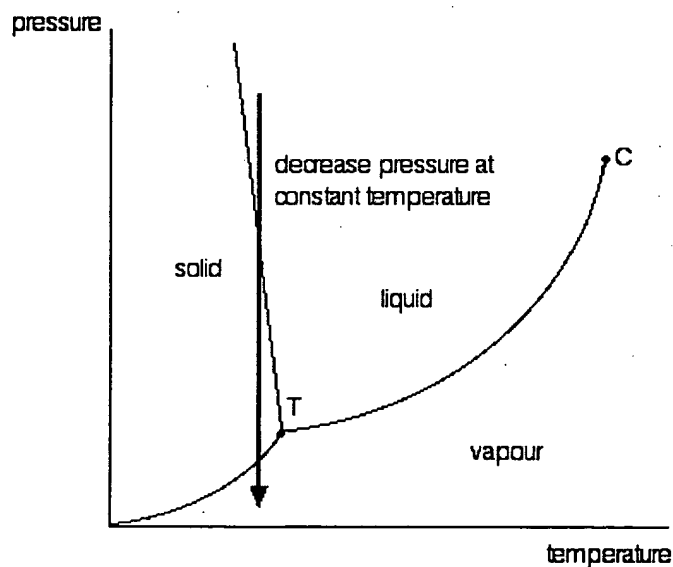
Notice that the triple point for water occurs at a very low pressure. Notice also that the critical temperature is 374°C . It would be impossible to convert water from a gas to a liquid by compressing it above this temperature.

The normal melting and boiling points of water are found in exactly the same way as we have already discussed - by seeing where the 1 atmosphere pressure line crosses the solid-liquid and then the liquid-vapour equilibrium lines.

Note: Further up the page I mentioned two ways in which these diagrams are distorted to make them easier to follow. I have already pointed out that the solid-liquid equilibrium line should really be much more vertical. This last diagram illustrates the other major distortion -

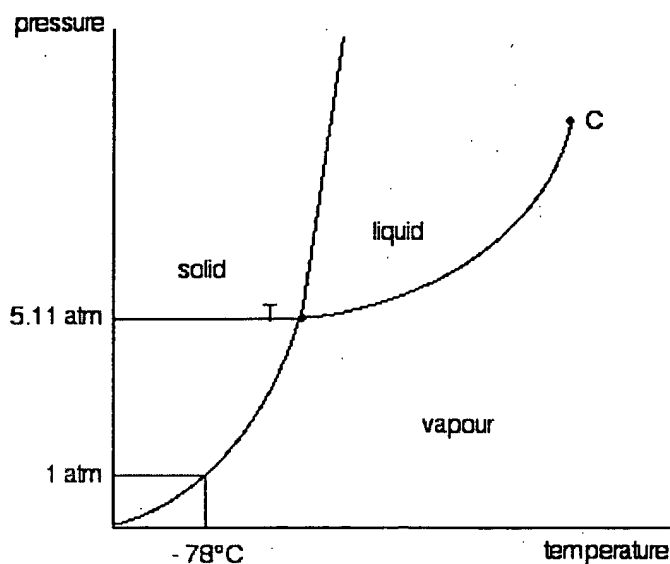
which is to the scales of both pressure and temperature. Look, for example, at the gaps between the various quoted pressure figures and then imagine that you had to plot those on a bit of graph paper! The temperature scale is equally haphazard.

Just one final example of using this diagram (because it appeals to me). Imagine lowering the pressure on liquid water along the line in the diagram below.



The phase diagram shows that the water would first freeze to form ice as it crossed into the solid area. When the pressure fell low enough, the ice would then sublime to give water vapour. In other words, the change is from liquid to solid to vapour. I find that satisfyingly bizarre!

The phase diagram for carbon dioxide



The only thing special about this phase diagram is the position of the triple point which is well above atmospheric pressure. It is impossible to get any liquid carbon dioxide at pressures less than 5.11 atmospheres.

That means that at 1 atmosphere pressure, carbon dioxide will sublime at a temperature of -78°C .

This is the reason that solid carbon dioxide is often known as "dry ice". You can't get liquid carbon dioxide under normal conditions - only the solid or the vapour.

Where would you like to go now?

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